Topic 12b – Redox
Revision Notes

1) Introduction

- Oxidation is loss of electrons
- Reduction is gain of electrons
- An oxidising agent accepts electrons (and is reduced in the process)
- A reducing agent donates electrons (and is oxidised in the process)

2) Oxidation states

- Oxidation states are "charges" assigned to each element in a reaction. The rules for assigning oxidation states are:
  - Elements are zero
  - In compounds, H is +1 and O is -2
  - In compounds, Group 1 elements are +1, Group 2 are +2, Group 6 are -2 and Group 7 are -1 (these are real charges)

- Be able to apply this to redox reactions of group 2 elements with water and oxygen e.g.

  $$\mathbf{Ba + 2H_2O \rightarrow Ba(OH)_2 + H_2}$$

  Ba: 0 to +2 = oxidation
  H: +1 to 0 = reduction

- The overall equation can be split into half-equations, one for the oxidation and one for the reduction:

$\mathbf{Ba \rightarrow Ba^{2+} + 2e^-}$	**oxidation (loss of electrons)**
$\mathbf{2H_2O + 2e^- \rightarrow 2OH^- + H_2}$	**reduction (gain of electrons)**

3) Constructing half-equations

- Constructing simple half-equations involves two steps: balancing atoms and balancing charge by adding electrons e.g. oxidation of iodide ions, I^-, to iodine, I_2.

Write down formulae:	$\mathbf{I^- \rightarrow I_2}$
Balance atoms:	$\mathbf{2I^- \rightarrow I_2}$
Balance charge:	$\mathbf{2I^- \rightarrow I_2 + 2e^-}$

- Constructing more complicated half-equations involves 2 extra steps, namely: balancing oxygens by adding water and balancing hydrogens by adding H^+ e.g. oxidation of V^{2+} to VO_3^-

Write down formulae:	$\mathbf{V^{2+} \rightarrow VO_3^-}$
Balance atoms (vanadiums):	not needed
Balance oxygens:	$\mathbf{V^{2+} + 3H_2O \rightarrow VO_3^-}$
Balance hydrogens:	$\mathbf{V^{2+} + 3H_2O \rightarrow VO_3^- + 6H^+}$ $(2+ \rightarrow 5+)$
Balance charge:	$\mathbf{V^{2+} + 3H_2O \rightarrow VO_3^- + 6H^+ + 3e^-}$

4) Redox equations

- To produce an overall redox equation, multiply one or both half-equations until the number of electrons is the same

  $$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$$
  $$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$$

- In this case, multiply the first half-equation by 5 to get 5 electrons in both.

  $$5Fe^{2+}(aq) \rightarrow 5Fe^{3+}(aq) + 5e^-$$

- Now add the half-equations together, cancelling the electrons at the same time.

  $$5Fe^{2+}(aq) + MnO_4^-(aq) + 8H^+(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$$

- Any species that appears on both sides of the equation needs to be cancelled e.g. H^+